The Weak Acid / Base Rxn

Recall a balanced equation.

What is on the left side?

The right side?

When you compare the concentration (molarity, [mol/L]) of all reactants and to all products you get something called and <u>equilibrium expression</u>.

Where “K” is the equilibrium constant for that rxn

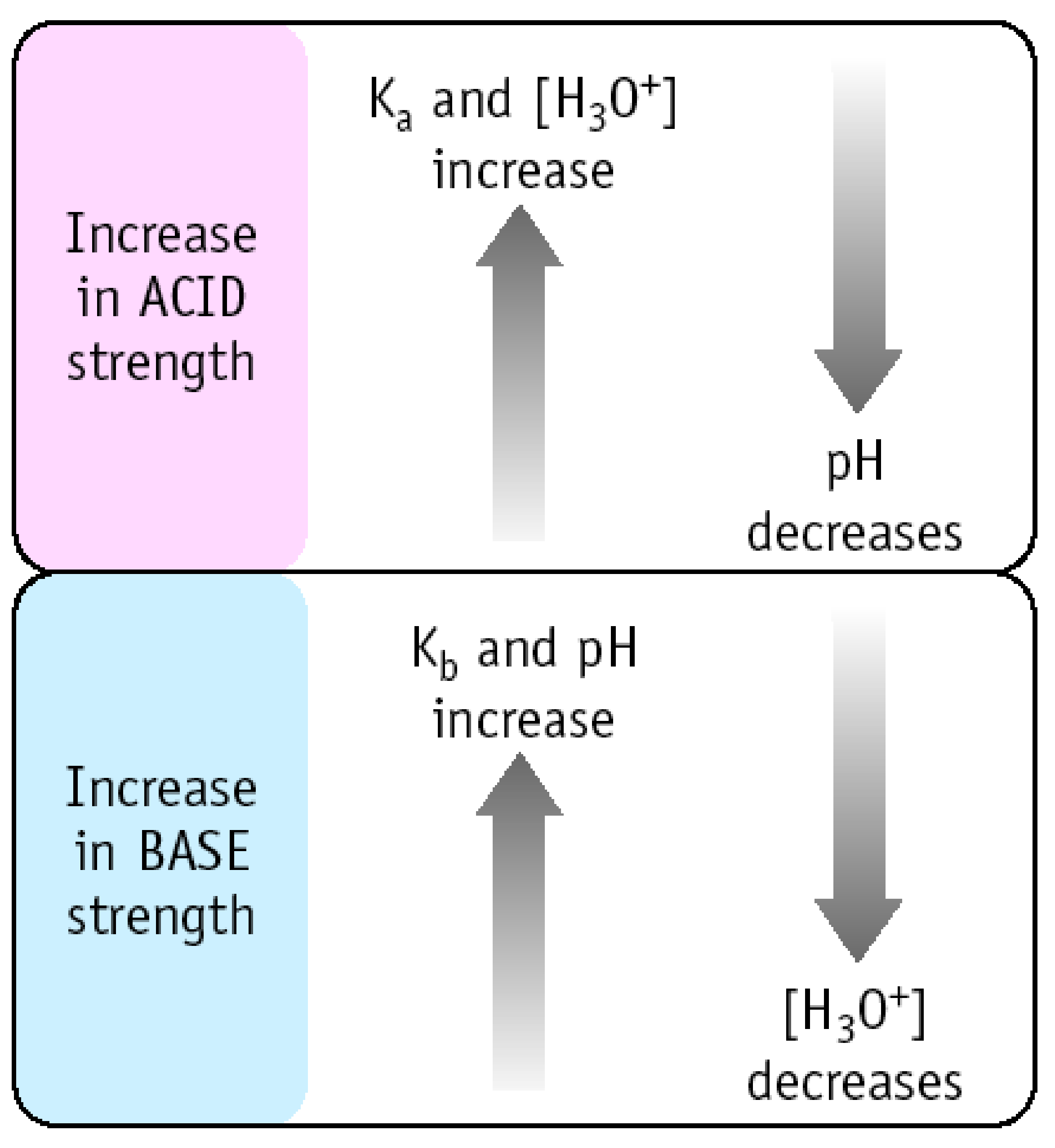
Equilibrium constants for Weak Acids

$$HA(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Weak acid has $K_a < 1$

Leads to small $[H_3O^+]$ and a pH of 2 - 7



Relation of K_a, K_b, $[H_3O^+]$ and pH

I.C.E. Table Method

I = Initial concentration, M

C = Change concentration, M

E = Equilibrium concentration, M

RXN	HA→	H^+	A^-
I	[HA]	0	0
C	-x	+x	+x
E	[HA]-x	x	x

I.C.E. Table Method

Step 1: Write balanced equation

Step 2: Write equilibrium (K_a) expression

Step 3: Construct ICE table and fill in

Step 4: Solve for unknown concentrations using equilibrium values, use $[H_3O^+]$ or $[H^+]$ to find pH if asked

Practice

If the K_a of an acid, “A,” is $1.5x10^{-2}$, what can you tell about its pH compared acid “B” who’s K_a is $6.2x10^{-4}$?

Equilibria involving weak acids

You have 1.00 M $HC_2H_3O_2$ Calc. the equilibrium concs. of $HC_2H_3O_2$, H^+, $C_2H_3O_2^{\ -}$, and the pH. ($K_a = 1.8x10^{-5}$)

Step 1. Write balanced equation.

$$HC_2H_3O_2 + H_2O \leftarrow\rightarrow H_3O^+ + C_2H_3O_2^{\ -}$$

or

$$HC_2H_3O_2 \leftarrow\rightarrow H^+ + C_2H_3O_2^{\ -}$$

Equilibria involving weak acids

You have 1.00 M $HC_2H_3O_2$ Calc. the equilibrium concs. of $HC_2H_3O_2$, H^+, $C_2H_3O_2^-$, and the pH. ($K_a = 1.8\times10^{-5}$)

Step 2. Write equilibrium (K_a) expression.

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Equilibria involving weak acids

You have 1.00 M $HC_2H_3O_2$ Calc. the equilibrium concs. of $HC_2H_3O_2$, H^+, $C_2H_3O_2^{\ -}$, and the pH. (K_a = 1.8×10^{-5})

Step 3. Construct ICE table and fill in.

	$[HC_2H_3O_2]$	$[H^+]$	$[C_2H_3O_2^{\ -}]$
Initial			
Change	1.00	0	0
Equilib.	-x	+x	+x
	1.00-x	x	x

Equilibria involving weak acids

You have 1.00 M $HC_2H_3O_2$ Calc. the equilibrium concs. of $HC_2H_3O_2$, H^+, $C_2H_3O_2^-$, and the pH. (K_a = 1.8×10^{-5})

Step 4. Solve for unknown concentrations using equilibrium values, use $[H_3O^+]$ or $[H^+]$ to find pH if asked

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = 1.8\times10^{-5} = \frac{[x^2]}{[1.0-x]}$$

This is a quadratic. Solve using quadratic formula.

or you can make an approximation if x is very small!

(Rule of thumb: 10^{-4} or smaller is ok)

Equilibria involving weak acids

You have 1.00 M $HC_2H_3O_2$ Calc. the equilibrium concs. of $HC_2H_3O_2$, H^+, $C_2H_3O_2^-$, and the pH. ($K_a = 1.8 \times 10^{-5}$)

Step 4. Solve K_a expression

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} \qquad K_a = \frac{[x][x]}{[1.0-x]}$$

First assume x is very small because K_a is so small.

$$K_a = 1.8 \times 10^{-5} = \frac{x^2}{1.00}$$

Equilibria involving weak acids

You have 1.00 M $HC_2H_3O_2$ Calc. the equilibrium concs. of $HC_2H_3O_2$, H^+, $C_2H_3O_2^-$, and the pH. (K_a = 1.8x10^{-5})

Step 4. Solve K_a approximate expression

$$K_a = 1.8 \times 10^{-5} = \frac{x^2}{1.00}$$

x = $[H^+]$ = $[C_2H_3O_2^-]$ = 4.2 x10^{-3} M

pH = - log$[H^+]$ = -log (4.2 x10^{-3})

= 2.37

Practice

What is the pH of a 0.020 M citric acid , $H_3C_6H_8O_7$, solution? $K_a = 7.5 \times 10^{-4}$

pH =? We need $[H^+]$ 1st $[H_3C_6H_8O_7] = 0.020M$

Weak acid, $K_a = 7.5 \times 10^{-4}$

Rxn	$H_3C_6H_8O_7$	H^+	$H_2C_6H_8O_7^-$
I			
C			
E			

$K_a =$ ____________

So since x = [H+] then pH must be?

Weak Acid Practice

If a 3.0M weak acid HF is used to etch glass, we need to find the concentration of $[H^+]$ ions in the solution.

1. Write the balance ionization/ dissociation of HF.
2. Write the K_a expression for HF ionization!

Practice

If a 3.0M weak acid HF is used to etch glass, we need to find the concentration of $[H^+]$ ions in the solution. Given the K_a of HF as 6.3×10^{-4}.

Recall part II

If a 3.0M weak acid HF is used to etch glass, we need to find the concentration of $[H^+]$ ions in the solution. Given the K_a of HF as 6.3×10^{-4}.

Now use the quadratic equation ☺

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Practice

What is the pH of a 0.020 M Sulfurous Acid, H_2SO_3, solution? $K_a = 1.4 \times 10^{-2}$

Rxn	H_2SO_3	H^+	HSO_3^-
I			
C			
E			

K_a is $>1.0\text{x}10^{-2}$ so we need to use quadratic to solve for "x"